THE INSTITUTE OF EDUCATION

Subject: Chemistry

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COURSE: Leaving Certificate

ACADEMIC LEVEL: Higher Level

ACADEMIC YEAR: 2021 - 2022

TOPIC: Shapes and Intermolecular Forces

CONTENTS

Shapes and Intermolecular Forces

Shapes of molecules

Remember: Covalent compounds exist as individual molecules - these molecules take on particular shapes according to the valence shell electron pair repulsion theory

Q: Describe the valence shell electron pair repulsion theory (V.S.E.P.R.T.)?

- This theory is used to **predict the shape of a molecule**
- The shape of the molecule depends on:
  - i) **Number of electron pairs** around the **central atom**
  - ii) The **type of these pairs** (bond pairs or lone pairs)

- The different **electron pairs repel each other** and try to be as far apart as possible
- **Two lone pairs** of electrons have the **strongest repulsion**, a **lone pair and bond pair** of electrons have a **weaker** repulsion, and **two bond pairs** of electrons have the **weakest repulsion**

YOU MUST LEARN THIS!

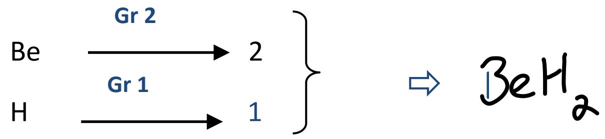
Electron pairs	Shape of molecule	Bond angle
Two bond pairs	Linear shape	180˚
Three bond pairs	Trigonal Planar shape	120˚
Four bond pairs	Tetrahedral shape	109.5˚
Three bond pairs and one lone pair	Pyramidal shape	107˚
Two bond pairs and two lone pairs	V- shaped/ bent	104.5˚

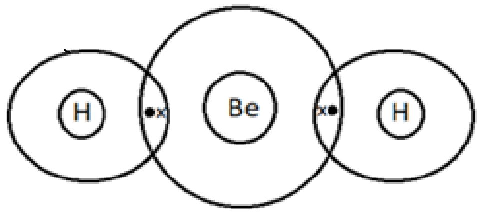
Remember:

All molecules with two atoms only are linear in shape – molecules with more than two atoms take on different shapes

ADD ABOUT THESE SHAPES APPLYING ONLY TO COVALENT MOLECULES - REMIND ABOIUT IONIC - COPY AND PASTE FROM CHEMCIAL BONDING CHAPTER

Example 1: Using the valence shell electron pair repulsion theory, predict the shape of a molecule of beryllium hydride and state the bond angle present in the molecule

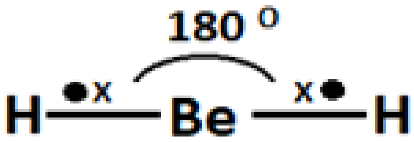


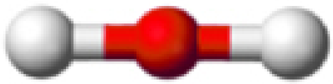


2 bond pairs of electrons

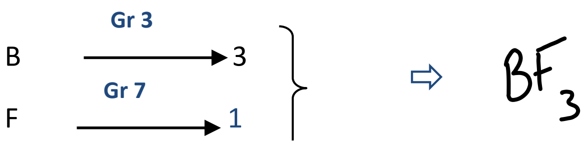
Linear Shape

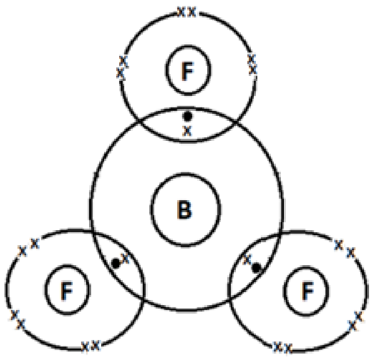
Bond angle - 180°





Example 2: Using the valence shell electron pair repulsion theory, predict the shape of a molecule of boron trifluoride and state the bond angle present

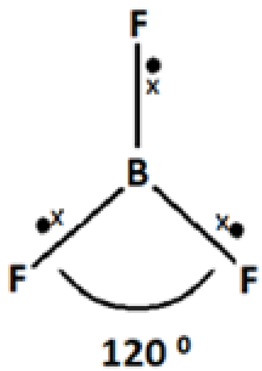


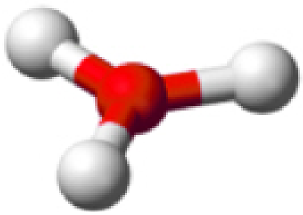


3 bond pairs of electrons

Trigonal Planar Shape

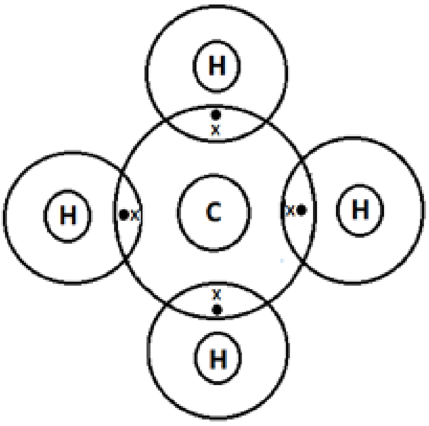
Bond angle - 120°





Example 3: Using the valence shell electron pair repulsion theory, predict the shape of a molecule of methane and state the bond angle present

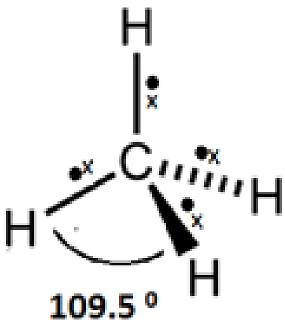
C —Gr 4→ 4
H —Gr 1→ 1
⇨ CH_4

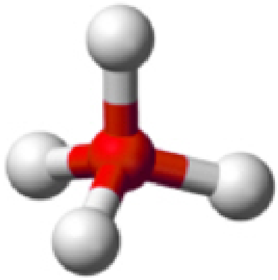


4 bond pairs of electrons

Tetrahedral Shape

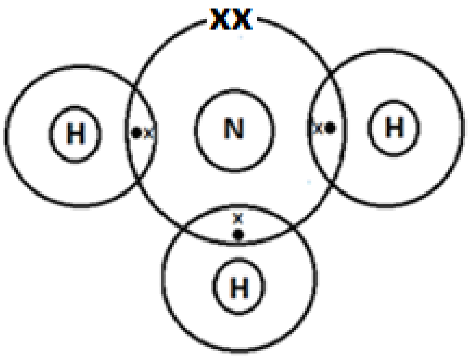
Bond angle - 109.5 º





Example 4: Using the valence shell electron pair repulsion theory, predict the shape of a molecule of ammonia and state the bond angle present

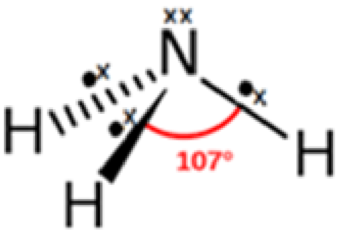
N —Gr 5→ 3
H —Gr 1→ 1
⇨ NH_3

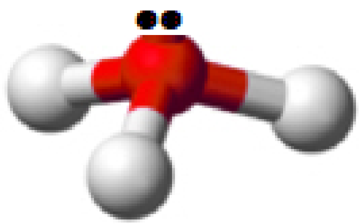


3 bond pairs of electrons
<u>and</u>
1 lone pair of electrons

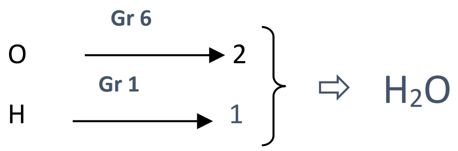
Pyrimidal Shape

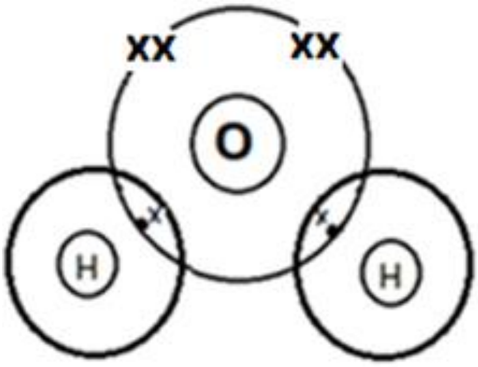
Bond angle - 107 º





Example 5: Using the valence shell electron pair repulsion theory, predict the shape of a molecule of water and state the bond angle present

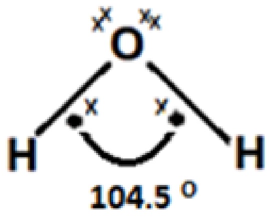


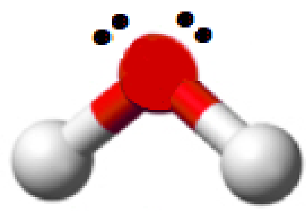


2 bond pairs of electrons
and
2 lone pair of electrons

V-Shaped

Bond angle - 104.5 °

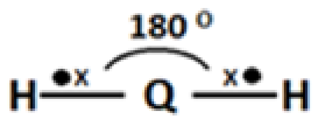




Example 6: Give the two possible shapes of a molecule with the general formula QH_2 where Q represents any element

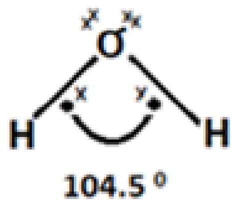
Note: In QH_2, there are three atoms where Q is the central atom and is bonded to two H atoms. This can result in two shapes:

Linear Shape



or

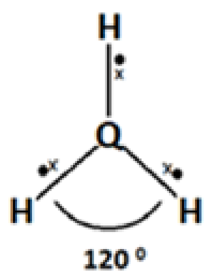
V-Shaped



Example 7: Give the two possible shapes of a molecule with the general formula QH_3 where Q represents any element

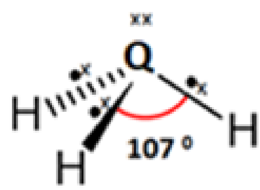
Note: In QH_3, there are four atoms where Q is the central atom and is bonded to three H atoms. This can result in two shapes:

Trigonal Planar Shape



or

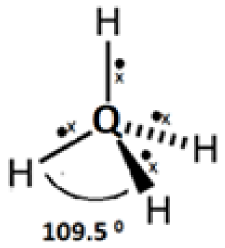
Pyramidal Shape



Example 8: Give the only possible shape of a molecule with the general formula QH_4 where Q represents any element

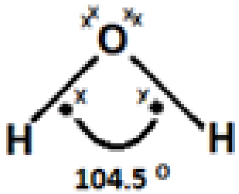
Note: In QH_4, there are five atoms where Q is the central atom and is bonded to four H atoms. This can only result in one shape:

Tetrahedral Shape

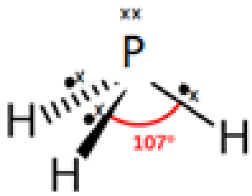


Q: Explain using the valence shell electron pair repulsion theory why water has a smaller bond angle than phosphine (phosphorus and hydrogen) but phosphine has a smaller bond angle than germane (germanium and hydrogen)

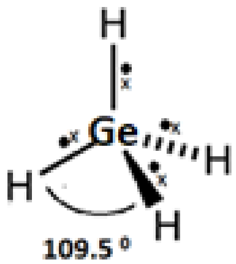
O —(Gr 6)→ 2
H —(Gr 1)→ 1
⇨ H_2O



P —(Gr 5)→ 3
H —(Gr 1)→ 1
⇨ PH_3



Ge —(Gr 4)→ 4
H —(Gr 1)→ 1
⇨ GeH_4



Water Vs Phosphine

- A **water** molecule has two bond pairs and **two lone pairs**, a **phosphine** molecule has three bond pairs and **one lone pair**
- **Lone pair - lone pair repulsion has a greater repelling power than lone pair - bond pair** repulsion and the water molecule bonds are pushed closer together (104.5 ˚) than the phosphine molecule bonds (107˚)

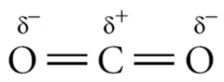
Phosphine Vs Germane

- A **phosphine** molecule has three bond pairs and **one lone pair**, a **germane** molecule has four bond pairs and **no lone pairs**
- **Lone pair - bond pair repulsion has a greater repelling power than bond pair - bond pair** repulsion and the phosphine molecule bonds are pushed closer together (107˚) than the germane molecule bonds (109.5˚)

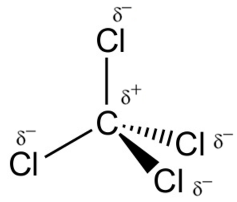
Q: Explain why certain molecules can be non-polar molecules despite having polar bonds within their molecule

- If the molecule is **highly symmetrical**
- The **centres of positive and negative charge coincide**
- The molecule as a whole is non-polar despite having polar bonds

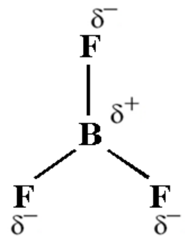
Examples of non-polar molecules with polar bonds:



Carbon Dioxide (CO_2)



Tetrachloromethane (CCl_4)



Boron trifluoride (BF_3)

Note: A molecule is highly symmetrical if:

1. It has a linear, trigonal planar or tetrahedral shape

and

2. The atoms coming from the central atom are the same element

Important:

- Polar and non-polar molecules display very different properties
  Example: Boiling points and solubilities;

- Care needs to be taken if the molecule is symmetrical as it will display non-polar properties

i.e. Have van der waals intermolecular forces between its molecules

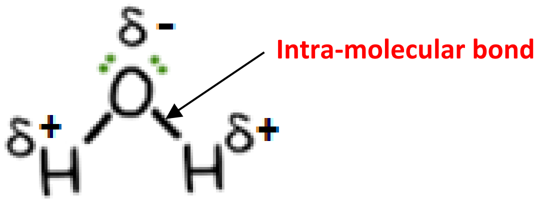
Typically have low boiling points

Will be completely insoluble in water

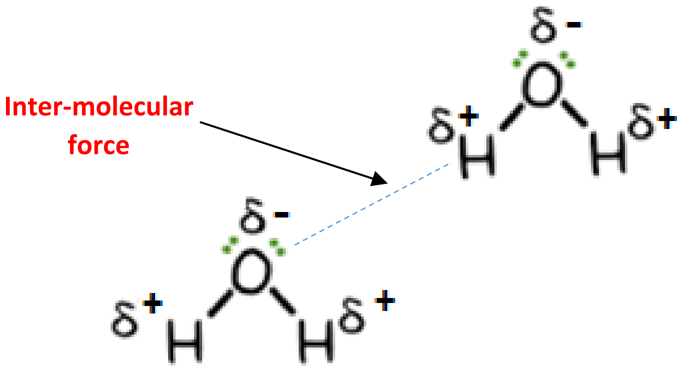
Inter-molecular forces

Q: Distinguish between intramolecular bonds and intermolecular forces

Intramolecular bonds - Type of **bond within** the **molecule**



Intermolecular forces – **Forces** of attraction that exist **between one molecule and another**



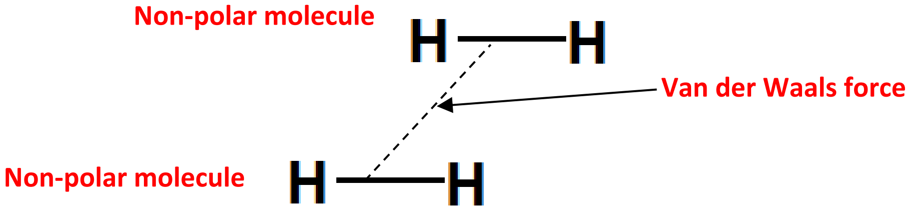
Q: Name and describe the three types of intermolecular forces

(1) Van der Waals forces

- Form between a **non-polar molecule and another non-polar molecule** – **caused by temporary dipoles** (partial charges) in the atoms due to movement of electrons
- Very **weak** forces - **do not require much energy to break**
- Therefore, molecules with van der Waals forces have **low melting and boiling points** – **Most are gases** at room temperature
- The strength of the **van der Waals force gets bigger** as the **molecule's molecular mass (size)** increases

symmetrical molecules

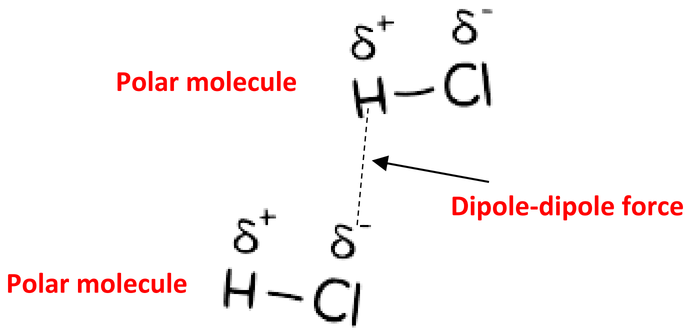
Examples: H_2, Cl_2, O_2, N_2, CH_4, **CO_2, CCl_4, BCl_3**



(2) Dipole-Dipole forces

- Form between **polar molecules – caused by permanent dipoles** in the atoms
- Stronger than van der Waals forces

Example: Hydrogen chloride,



Notice: The **slightly positive atom** in one molecule is **attracted** to the **slightly negative atom** in another molecule

(3) Hydrogen Bonding (H-Bonding)

- Form between **specific polar molecules**
- **Strongest** intermolecular forces

Q: What is a hydrogen bond (H-Bond)?

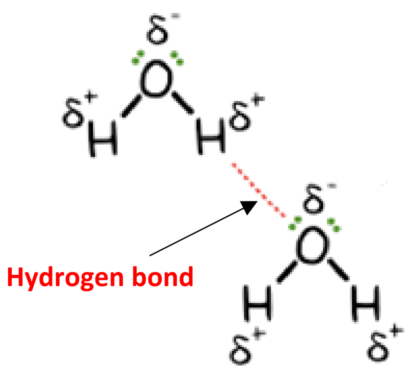
- A hydrogen bond is the **attraction between a slightly positive hydrogen (H)** in one molecule and a slightly negative **nitrogen (N), oxygen (O) or fluorine (F)** in another molecule

Q: Why does hydrogen bonding occur between molecules with slightly negative nitrogen, oxygen, or fluorine atoms?

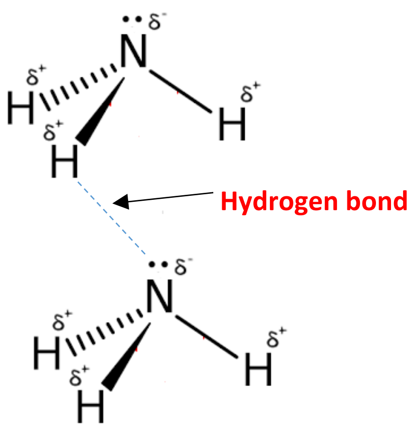
- Nitrogen, oxygen, and fluorine are **small** atoms and **highly electronegative**

Examples:

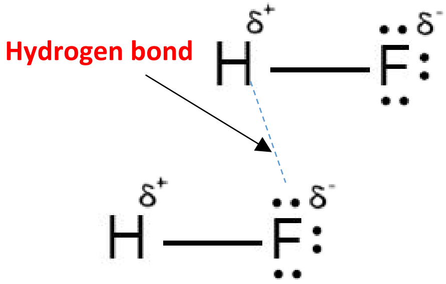
Water



Ammonia



Hydrogen fluoride



Comparing boiling points of substances

Important: When a substance boils, the molecules receive enough energy for the intermolecular forces between them to break. The molecules can then move away independently

ALWAYS: When comparing boiling points of different substances check:

1st - type of intermolecular force …….hydrogen bonds > dipole-dipole > Van der Waals

2nd - the molecular mass of the molecule……….heavier molecules have higher boiling points

3rd – how polar the molecule is……..more polar molecules have higher boiling points

4th – how many intermolecular forces form between the molecules…… more intermolecular forces have higher boiling points

Example 1: The boiling point of water is 100 ˚C while the boiling point of hydrogen sulfide is only - 61 ˚C. Account for the difference in boiling point between these substances.

- **Hydrogen sulfide** molecules are non-polar – have **weak van der Waals forces** between their molecules
- **Water** molecules are polar and have **very strong hydrogen bonds** between their molecules (have slightly positive hydrogen atoms attracted to slightly negative oxygen atoms)
- More energy is required to break the forces in water, explaining the difference in boiling points

Example 2: Explain why hydrogen chloride only has a boiling point of -85 ˚C while methane has a boiling point of -161.5 ˚C

- **Methane** molecules are **non-polar** – have **weak van der Waals forces** between their molecules
- **Hydrogen chloride** molecules are **polar** and have **stronger dipole-dipole forces** between their molecules
- More energy is required to break the forces in hydrogen chloride, explaining its higher boiling point

Example 3: The boiling points of hydrogen and oxygen are 20.0 K and 90.2 K respectively. Account for the higher boiling point of oxygen

- Both hydrogen and oxygen molecules are non-polar – **both have weak van der Waals forces** between the molecules
- However, **oxygen** (32) **has a larger molecular mass** than hydrogen (2) so has **stronger intermolecular forces** between its molecules and more energy is required to break them

Example 4: Suggest a reason why the boiling point of ammonia (-33˚C) is significantly lower than that of water (100˚C)

Ammonia (NH_3) N: 3.04
H: 2.20

Difference = 0.84 ....less than 1.7
greater than 0.4 ....Polar covalent bond

Water (H_2O) O: 3.44
H: 2.20

Difference = 1.24 ....less than 1.7
greater than 0.4 ....Polar covalent bond

- Both molecules are polar and have strong hydrogen bonds between their molecules and have very similar molecular masses
- However, as the **N-H bond in ammonia is less polar** than the O-H bond in water, **weaker hydrogen bonds** form between ammonia's molecules
- **Less energy** is **required to break these bonds**, which explains ammonia's lower boiling point

Example 5: Suggest why water (100˚C) has a higher boiling point than hydrogen fluoride (19.5 ˚C).

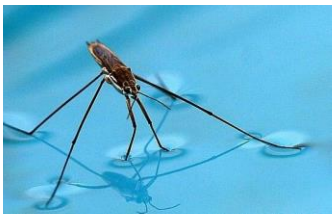
- Both molecules are polar and have strong hydrogen bonds between their molecules and have similar molecular masses
- An individual hydrogen bond in hydrogen fluoride is stronger than an individual hydrogen bond in water as the H-F bond in hydrogen fluoride is more polar than the O-H bond in water
- However, **more hydrogen bonds form around** a **water** molecule than around a hydrogen fluoride molecule
- More energy is required to break more bonds, which explains water's higher boiling point

Q: Give three effects of hydrogen bonding?

i. It explains why the three hydrides **H_2O, HF and NH_3** have **higher boiling points** than **other hydrides**

ii. It explains why **H_2O** has a **higher boiling point** than other **molecules** of **similar molecular mass**

iii. It explains the **surface tension ('skin effect')** on water – The water molecules on the surface are strongly attracted to each other by hydrogen bonding

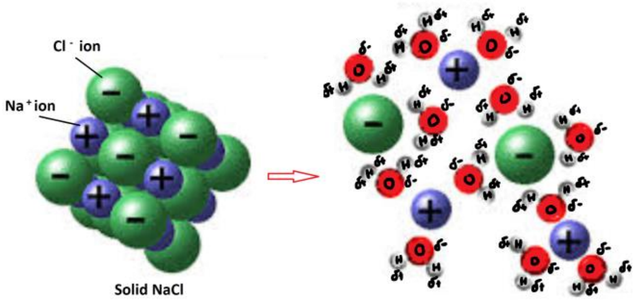


Predicting solubilities of substances

1) Ionic compounds **Example:** Sodium chloride (NaCl)

a) In water i.e. a polar solvent

- Ionic compounds are **very soluble in water** –
  - The **partial negative charge** in the polar water is **attracted to** the **positive ion** in the ionic compound
  - The **partial positive charge** in the polar water is **attracted to** the **negative ion** in the ionic compound
- The **crystalline structure** can be **pulled apart** and the ionic compound dissolves



Water molecule

Note: Some ionic compounds are not soluble in water at room temperature

Examples: Silver Chloride (AgCl)
Barium sulfate ($BaSO_4$)
Magnesium carbonate ($MgCO_3$)
Calcium carbonate ($CaCO_3$)

b) In a non-polar solvent **Example: Cyclohexane**

- Ionic compounds are **not soluble** in non-polar solvents
- The are no partial charges in the non-polar molecule to pull the ionic crystal apart

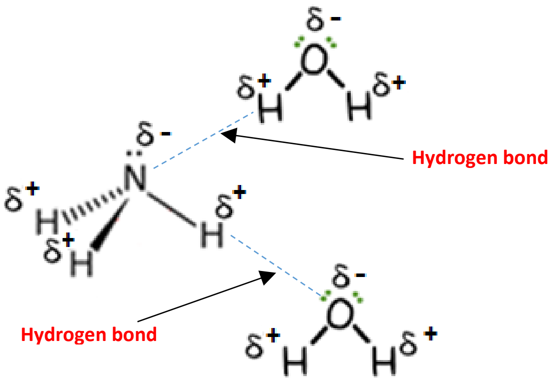
2) Covalent compounds

Rule: Polar dissolves polar

Non-polar dissolves non-polar

Q: Explain why ammonia is readily soluble in water? Use a labelled diagram to aid your answer

- Water molecules are **polar**, ammonia molecules are **polar**
- Attractive forces - **hydrogen bonds** form between the molecules and the **ammonia dissolves**

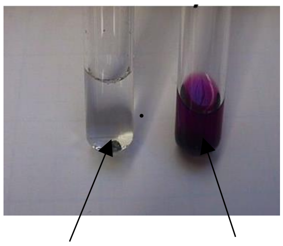


Q: Would you expect iodine to be soluble in water? Explain

- Iodine is not soluble/sparingly soluble in water
- Iodine molecules are **non-polar**, water molecules are **polar**
- **Little or no attractive forces form** between the molecules, so the iodine does not dissolve

Q: Name an organic solvent iodine is readily soluble in. Explain why this is the case

- **Cyclohexane** (C_6H_{12})
- Iodine molecules are **non-polar**, cyclohexane molecules are **non-polar**
- **Attractive forces form** between the molecules and the iodine dissolves



Exam Question: Consider the data shown below for a number of simple hydrides.

Hydride formula	Hydride name	Boiling point (°C)
CH_4	methane	−161.5
H_2O	water	100.0
H_2S	hydrogen sulfide	−60.0
H_2Se	hydrogen selenide	−41.3
H_2Te	hydrogen telluride	−2.2

(*i*) Why is water the only hydride in the table that exhibits strong hydrogen bonding?

(*ii*) Why is the boiling point of hydrogen selenide higher than that of hydrogen sulfide?

(*iii*) Do you expect hydrogen sulfide to be water soluble? Explain

(iv) Explain why, although carbon and selenium have the same electronegativity value, **CH_4** is almost completely insoluble in water and **H_2Se** is slightly soluble in water.

(i)

- The intermolecular forces in **water** involve an attraction between a slightly positive hydrogen in one water molecule and a **slightly negative oxygen** in another water molecule.
- **Nitrogen, oxygen and fluorine are small electronegative atoms** **required for hydrogen bonding** to occur
- The other hydrides in the table do not possess these elements

(ii)

- Both hydrogen selenide and hydrogen sulfide molecules are non-polar – **both have weak van der Waals forces** between the molecules
- However, hydrogen selenide (81) **has a larger molecular mass** than hydrogen sulfide (34) so has **stronger intermolecular forces** between its molecules are and more energy is required to break them

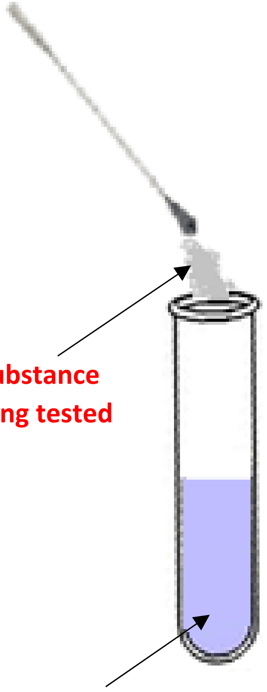
(iii)

- No hydrogen sulfide is not soluble/sparingly soluble in water
- Hydrogen sulfide molecules are non-polar/very weakly polar, water molecules are polar
- Little or no attractive forces form between the molecules and the hydrogen sulfide does not dissolve easily

(iv)

- Methane is a **perfectly symmetrical molecule**
- The **centres of positive and negative charge coincide**
- The molecule is **completely non-polar** and will be **completely insoluble** in polar water
- Hydrogen selenide can be described as weakly polar and will show some slight solubility in water – it is not a highly symmetrical molecule

Testing solubility in different solvents of ionic and covalent substances



- Half fill 5 test tubes with **water**
- Add the following substances, one to each test tube

  i. Sodium chloride
  ii. Iodine
  iii. Glucose
  iv. Cooking oil
  v. Ethanol

  Note: Use a spatula for solids and a dropper for liquids

- Stopper and shake the test tube well and observe if the substance has dissolved

- Half fill 5 new test tubes with **cyclohexane** and repeat the solubility test for the same five substances

Results:

Substance tested	Type of bonding	Soluble in water?	Soluble in cyclohexane?
Sodium Chloride	Ionic	Yes	No
Iodine	Non-polar covalent	No	Yes
Glucose	Polar covalent	Yes	No
Cooking oil	Non-polar covalent	No	Yes
Ethanol	Polar covalent	Yes	No

Shapes, Inter-molecular forces, and solubility - Questions

Q1:

a) Outline the valance shell electron pair repulsion theory

b) State the shape and bond angle given to a molecule if it has
   - i) two bond pairs
   - ii) three bond pairs
   - iii) four bond pairs
   - iv) three bond pairs and one lone pair
   - v) two bond pairs and two lone pairs

Q2:

a) *Use the valence shell electron pair repulsion theory to predict the shape and state the bond angle in:

i) Boron trichloride

ii) Water

iii) Germane (GeH_4)

iv) Ammonia

v) Beryllium hydride

vi) Arsine (AsH_3)

vii) Methane

viii) Phosphine

ix) Hydrogen sulfide

x) Oxygen Difluoride

xi) Silane (SiH_4)

xii) Carbon dioxide

b) Give the two possible shapes a molecule with general formula AB_2 could form

c) Give the two possible shapes a molecule with general formula AB_3 could form

d) Give the one possible shape a molecule with general formula AB_4 can form

Q3:

a) Distinguish between intra-molecular bonds and intermolecular forces

b) Name the three types of intermolecular forces

c) Between what types of molecules do van der Waals forces form?

d) Explain why molecules with van der Waals forces between them have low melting and boiling points

e) Explain why oxygen has a higher boiling point than hydrogen

f) **Between what types of molecules do dipole-dipole forces form?**
g) **Explain why molecules with dipole-dipole forces have higher melting and boiling points than molecules with van der waals forces between their molecules**
h) **Draw a labelled diagram to show the dipole-dipole forces between two molecules of hydrogen chloride**
i) **What is a hydrogen bond?**
j) **Give <u>two</u> properties of oxygen, nitrogen and fluorine atoms that results in them forming hydrogen bonds with hydrogen atoms.**
k) **Explain why molecules with hydrogen bonding between their molecules have higher melting and boiling points than molecules with van der waals forces or dipole-dipole forces between their molecules**
l) **Draw a labelled diagram to show the hydrogen bonding between two molecules of**
   i. **Water**
   ii. **Ammonia**
   iii. **Hydrogen fluoride**

Q4:

a) **Use your knowledge of intermolecular forces to explain why methane has a very low boiling point (b.p. = - 164 ˚C)**
b) **The boiling points of hydrogen gas, hydrogen chloride and hydrogen fluoride are253°C, - 85°C and 19.5˚C. Account for the different boiling points in these three molecules.**
c) **Explain why chlorine has a boiling point of - 34˚C but fluorine has a boiling point of -188˚C.**
d) **Explain why hydrogen chloride has a higher boiling point than oxygen**
e) **Hydrogen sulfide has a boiling point of 212.3 K and water has a boiling point of 373 K. Account for the difference in the boiling point of these substances.**
f) **Account for the higher boiling point of water compared to ammonia**
g) **Account for the higher boiling point of water compared to hydrogen fluoride.**
h) **Give three effects explained by hydrogen bonding**

Q5: Explain why CO_2, BF_3, BCl_3, CCl_4 are non-polar molecules despite having polar bonds within their molecules

***Q6: State the type of intermolecular forces that form between the following molecules**

(Be careful if molecule is symmetrical!)

a) **Phosphine (PH_3)**
b) **Water**
c) **Methane (CH_4)**
d) **Ammonia (NH_3)**
e) **Arsine (AsH_3)**
f) **Carbon dioxide**
g) **Hydrogen chloride**
h) **Hydrogen fluoride**
i) **Hydrogen sulfide**
j) **Boron trifluoride**

Q7:

a) **Comment on the solubility of ionic compounds in water**
b) **Name four ionic compounds that are not soluble in water**
c) **Comment on the solubility of ionic compounds in non-polar solvents**
d) **Explain with the aid of a labelled diagram why ammonia is very soluble in water?**
e) **Explain why iodine is not soluble in water**
f) **Name a solvent iodine is soluble in and explain why iodine is soluble in this solvent**
g) **Would you expect hydrogen sulfide to be soluble in water? Explain your answer**

***Q8:** Consider the three gaseous hydrides **NH_3**, **PH_3** and **AsH_3**.

Molecular formula	Common name	IUPAC name	Boiling point (°C)
NH_3	ammonia	azane	– 33.3
PH_3	phosphine	phosphane	– 87.7
AsH_3	arsine	arsane	– 62.5

Using data in the *Formulae and Tables* booklet determine the type of bonding arsine is expected to have.

Draw a dot and cross diagram (valence electrons are sufficient) to show the bonding in arsine.
Predict the shape of the arsine molecule.

Which, if any, of the three hydrides would you expect to have intermolecular hydrogen bonding?
Justify your answer.

Suggest a reason why
(*i*) ammonia has the highest boiling point of the three hydrides,
(*ii*) phosphine's boiling point is lower than that of arsine.

Q9: A student wanted to investigate the solubility of various substances in a polar solvent and nonpolar solvent. The substances investigated were 1) Calcium fluoride, 2) Pentane (C_5H_{12}), 3) Glucose, 4) Iodine, and 5) Ethanol.

a) **Name a suitable polar solvent and suitable nonpolar solvent for this investigation**
b) **Describe how the investigation is carried out**
c) **In your answer book rewrite and complete the table of the substances investigated**

Substance tested	Type of bonding	Soluble in polar solvent?	Soluble in nonpolar solvent?
1) Calcium fluoride			
2) Pentane (C_5H_{12})			
3) Glucose			
4) Iodine			
5) Ethanol			

Calculation/Working out answers

Q2:

a)

i)	**Shape: Trigonal planar**	**Bond angle: 120°**
ii)	**Shape: V-Shaped**	**Bond angle: 104.5°**
iii)	**Shape: Tetrahedral**	**Bond angle: 109.5°**
iv)	**Shape: Pyramidal**	**Bond angle: 107°**
v)	**Shape: Linear**	**Bond angle: 180°**
vi)	**Shape: Pyramidal**	**Bond angle: 107°**
vii)	**Shape: Tetrahedral**	**Bond angle: 109.5°**
viii)	**Shape: Pyramidal**	**Bond angle: 107°**
ix)	**Shape: V-Shaped**	**Bond angle: 104.5°**
x)	**Shape: Tetrahedral**	**Bond angle: 109.5°**
xi)	**Shape: Linear**	**Bond angle: 180°**

b) Answer: Shape 1: Linear- Bond angle: 180°; Shape 2: V-Shaped- Bond angle: 104.5°

c) Answer: Shape 1: Trigonal planar- Bond angle: 120°; Shape 2: Pyramidal- Bond angle: 107°

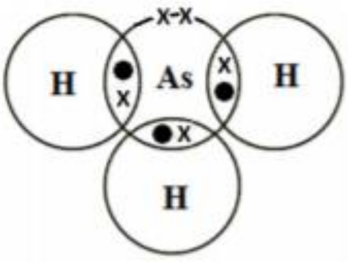
Q6:

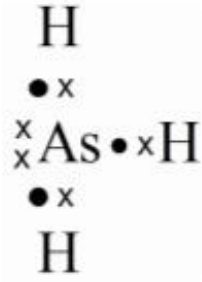
a) **Van der Waals**
b) **Hydrogen bonding**
c) **Van der Waals**
d) **Hydrogen bonding**
e) **Van der Waals**
f) **Van der Waals**
g) **Dipole- Dipole**
h) **Hydrogen bonding**
i) **Dipole- Dipole**
j) **Van der Waals**

Q8:

DETERMINE: **pure covalent / non-polar** covalent **(4)**
[Accept **slightly polar** covalent.]

DRAW:





(6)

PREDICT: **pyramidal (trigonal pyramidal)** (3)
[No marks for 'trigonal' on its own.]

[Diagram insufficient unless a 3d representation using at least one wedge or dash, as shown.]

WHICH: **azane / ammonia / NH_3** (3)

JUSTIFY: **nitrogen (N) small and** strongly **electronegative** / hydrogen bonds occur when **hydrogen is bonded to nitrogen**, oxygen, fluorine **(N,** O, F) (3)
[JUSTIFY marks only available if WHICH marks awarded.]

SUGGEST: (*i*) **hydrogen bonding** in azane (ammonia, NH_3) **stronger** than the other intermolecular forces (dipole-dipole, van der Waals) in phosphane or arsane / **hydrogen bonding strongest** type of intermolecular force // (3)

(*ii*) **phosphine (PH_3) molecules smaller (lighter, smallest** of the two) **/ phosphine (PH_3) molecules have smaller (smallest** of the two) **molecular mass / phosphine (PH_3) has fewer (fewest** of the two) **electrons / phosphine (PH_3) has a smaller (smallest** of the two) degree of **intermolecular (van der Waals) forces** (3)
[Accept equivalent phrases with respect to arsine (AsH_3)]